**PART 1**

1&2. A 0.1 M solution of an electrolyte has a pH of 4. The electrolyte is a ______________.

(a) strong acid  (b) strong base  (c) weak acid  (d) weak base  (e) neutral

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3&4. Which of the following salts has the lowest molar solubility?

(a) PbCl₂  (b) CaF₂  (c) PbBr₂  (d) BaF₂  (e) MgF₂

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5&6. CaSO₄ would be least soluble in

(a) 0.10 M Ca(NO₃)₂  
(b) 0.10 M Na₂SO₄  
(c) 0.10 M CaCl₂  
(d) 0.10 M K₂SO₄  
(e) It is equally soluble in all of these solutions.
7&8. Which of the following solutions is a buffer?

(a) 0.1 M HClO₃ and 0.2 M NaClO₃
(b) 0.1 M NaOH and 0.1 M NaF
(c) 0.1 M HCl and 0.2 M NaCl
(d) 0.1 M HCN and 0.2 M NaCN
(e) 0.1 M NH₃ and 0.1 M NaCl

9&10. What is the approximate pH of a solution that is 0.20 M in HNO₂ and 0.40 M in KNO₂?

(a) 3.65  (b) 3.22  (c) 2.83  (d) 3.45  (e) 3.35

11&12. One liter of 0.020 M NaI and one liter of 2.0 \times 10^{-6} \text{ M Pb(NO}_3\text{)}_2 are mixed. What is the value of Q_{sp} for PbI₂ in the final solution and will a precipitate form?

(a) Q_{sp} = 4.0 \times 10^{-8}; yes, a precipitate will form
(b) Q_{sp} = 8.0 \times 10^{-9}; no, a precipitate will not form
(c) Q_{sp} = 1.0 \times 10^{-8}; no, a precipitate will not form
(d) Q_{sp} = 4.0 \times 10^{-10}; yes, a precipitate will form
(e) Q_{sp} = 1.0 \times 10^{-10}; no, a precipitate will not form
13&14. The pH of a 0.200 M solution of an unknown weak monoprotic acid is 3.25. What is the $K_a$ of the unknown acid?

(a) $1.6 \times 10^{-6}$  (b) $2.9 \times 10^{-6}$  (c) $1.5 \times 10^{-5}$  (d) $4.0 \times 10^{-6}$  (e) $7.9 \times 10^{-3}$

15&16. Which of the following acid solutions (all at 0.10 M) has the highest concentration of anion?

(a) CH$_3$COOH
(b) CHOOH
(c) HNO$_2$
(d) HClO
(e) HCN
17&18. Suppose you have 100 mL of a 0.0010 M Cu(NO₃)₂ solution. If solid sodium hydroxide, NaOH, is slowly added to the beaker, at what pH would precipitation begin?

(a) 6.10  (b) 6.54  (c) 7.00  (d) 7.79  (e) 8.97

19&20. What is the pH of 0.45 M NaCH₃COO?

(a) 7.00  (b) 4.31  (c) 4.86  (d) 9.19  (e) 9.30
The following 5 questions deal with a single titration:

21&22. A 100.0 mL sample of 0.300 M methylamine, CH₃NH₂, is titrated with 0.200 M HCl. Calculate the initial pH before the titration is begun.

(a) 11.05  (b) 11.21  (c) 11.87  (d) 12.35  (e) 12.09

23&24. A 100.0 mL sample of 0.300 M methylamine, CH₃NH₂, is titrated with 0.200 M HCl. Calculate the pH after 50.0 mL of 0.200 M HCl has been added.

(a) 11.61  (b) 11.00  (c) 11.93  (d) 11.37  (e) 10.84
(5 pts) **25.** A 100.0 mL sample of 0.300 M methylamine, CH₃NH₂, is titrated with 0.200 M HCl. Calculate the pH at the equivalence point.

Will the solution be ACIDIC, BASIC, or NEUTRAL? (Circle the correct answer)

(5 pts) **26.** A 100.0 mL sample of 0.300 M methylamine, CH₃NH₂, is titrated with 0.200 M HCl. Calculate the pH after 175 mL of 0.200 M HCl is added.
27. A 100.0 mL sample of 0.300 M methylamine, CH₃NH₂, is titrated with 0.200 M HCl. Using the answers to Questions 21-26, sketch the titration curve with pH on the vertical axis and milliliters of acid added on the horizontal axis. Label the axes and plot your 4 points. Point out the buffer region and the equivalence point. If you cannot complete the calculations, sketch what the curve should look like for partial credit.

28. A buffer is prepared by mixing 1.00 mole of HCN and 2.00 moles of NaCN in 1.00 liter solution. To 200. mL of this solution is added 15.0 mL of 4.00 M HCl. What is the pH of the resulting solution?
(3 pts) **29.** (a) Explain what is happening in the beaker when a saturated solution of Fe(OH)$_3$ is prepared by adding solid iron(III) hydroxide to pure water. Draw what is happening in solution to illustrate your point.

(b) Write the equilibrium and the $K_{sp}$ expression for this system.

(c) Define molar solubility.

(d) Calculate the molar solubility, $s$, at 25°C, for Fe(OH)$_3$.

(5 pts) **30.** What is the pH of a solution of KCN? (<7, =7, or >7)? Explain using an appropriate equilibrium.